UNIT J

Reactions in Solutions

Types of Reactions in Solution

Three common types of reactions in solution are single replacement reactions, precipitation reactions and neutralizing reactions. The latter two represent different types of double replacement reactions.

1. Single Replacement Reactions

Single replacement reactions usually take place when a sample of an element is added to an aqueous solution of a compound. The element reacts with a compound to form a new element and a new compound. The general equation for a single replacement reaction is

$$A + BC \longrightarrow B + AC$$

If A is a metal, it will replace B to form AC; if A is a nonmetal, it will replace C to form BA. Some examples of these types of reactions are:

a) active metal + acid ⟶ hydrogen + a *salt*

$$Zn_{(s)} + 2HC\ell_{(aq)} \longrightarrow H_{2(g)} + ZnC\ell_{2(aq)}$$

$$2A\ell_{(s)} + 3H_2SO_{4(aq)} \longrightarrow 3H_{2(g)} + A\ell_2(SO_4)_{3(aq)}$$

Substances referred to as *salts* are ionic compounds consisting of a metal ion (or ammonium ion) combined with a simple or complex anion.

b) active metal + water ⟶ hydrogen + *base* or metal oxide

$$2Na_{(s)} + 2H_2O \longrightarrow H_{2(g)} + 2\ NaOH_{(aq)}$$

$$Ca_{(s)} + 2H_2O \longrightarrow H_{2(g)} + Ca(OH)_{2(aq)}$$

$$3\ Fe_{(s)} + \underset{\text{steam}}{4H_2O} \longrightarrow 4H_{2(g)} + Fe_3O_{4(aq)}$$

Substances referred to as *bases* are ionic compounds consisting of a metal ion (or ammonium ion) combined with a hydroxide ion, OH^-.

c) metal + salt ⟶ metal + salt

$$Zn_{(s)} + CuSO_{4(aq)} \longrightarrow Cu_{(s)} + ZnSO_{4(aq)}$$

d) nonemtal + salt ⟶ nonemtal + salt

$$C\ell_{2(g)} + 2NaBr_{(aq)} \longrightarrow Br_{2(\ell)} + 2NaC\ell_{(aq)}$$

$$Br_{2(\ell)} + 2KI_{(aq)} \longrightarrow I_{2(s)} + 2KBr_{(aq)}$$

For single replacement reactions involving a metal, the starting free metal must be more "reactive" than the metal ion in the salt that it displaces. Similarly, for single replacement reactions involving a nonmetal, the starting free nonmetal must be more "reactive" than the nonmetal ion in the salt that it displaces. However, for purposes of this course, assume that all the single replacement reactions considered will occur as given.

2. Double Replacement Reactions

Double replacement reactions usually occur when an aqueous solution of one compound is added to an aqueous solution of another compound. The two compounds react with each other to produce two different compounds. The general form of the equation for double

replacement reactions is

$$AB + CD \longrightarrow AD + CB$$

Double replacement reactions may be considered as involving an exchange of positive and negative species, where the positive species (A) of the first compound combines with the negative species (D) of the second compound, and the positive species (C) of the second compound combines with the negative species (B) of the first compound. In writing the formulas of the products, the charges of the combining species must be considered. Two common types of double replacement reactions are precipitation reactions and neutralization reactions.

a. Precipitation Reactions

Precipitation reactions occur when two compounds in solution react to produce a new compound that has low solubility. The new compound of low solubility is called a precipitate. Precipitation is identified in an equation by a product with the subscript "s". For example:

$$BaC\ell_{2(aq)} + H_2SO_{4(aq)} \longrightarrow 2HC\ell_{(aq)} + BaSO_{4(s)}$$

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow 2KNO_{3(aq)} + PbI_{2(s)}$$

Solid particles of precipitate settle out and can be separated from the remaining solution by filtration. The liquid which goes through the filter paper is called the *filtrate*. The solid precipitate caught in the filter paper can be dried and its mass determined.

The following groups of compounds are always soluble in water and therefore do not precipitate:

1. any compound of an alkali metal ion (Li^+, Na^+, K^+, Cs^+ or Rb^+)
2. any compound containing the ammonium ion (NH_4^+)
3. any inorganic acid
4. any compound containing the nitrate ion (NO_3^-)
5. any compound containing the acetate ion (CH_3COO^-)

For the solubility of compounds other than those mentioned above, refer to a table of solubilities.

These simple rules can be used to predict a precipitate in many reactions. For example, the reaction, $BaC\ell_{2(aq)} + K_2CrO_{4(aq)} \longrightarrow 2KC\ell_{(aq)} + BaCrO_{4(s)}$

produces a yellow precipitate. The $KC\ell$ is soluble since all K^+ compounds are soluble, therefore, the precipitate must be $BaCrO_4$.
As a second example, consider the reaction,

$$Co(NO_3)_{2(aq)} + 2NaOH_{(aq)} \longrightarrow Co(OH)_{2(s)} + 2NaNO_{3(aq)},$$ which

produces a blue precipitate. The precipitate cannot be $NaNO_3$ since all Na^+ compounds and all NO_3- compounds are soluble, therefore, the precipitate must be $Co(OH)_2$.

b. Neutralization Reactions

Neutralization reactions are reactions between acids and bases in solution. The products of a neutralization are water and salt.

$$\text{acid} + \text{base} \longrightarrow \text{water} + \text{a salt}$$

For example:

$$HC\ell_{(aq)} + LiOH_{(aq)} \longrightarrow HOH_{(\ell)} + LiC\ell_{(aq)}$$

$$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow 2HOH_{(\ell)} + Na_2SO_{4(aq)}$$

Acids and Bases - Some General Properties

All acids have the following properties in common:

a. form electrolytic solutions
b. taste sour
c. turn blue litmus red
d. react with zinc to form hydrogen gas
e. neutralize bases

All bases have the following properties in common:

a. form electrolytic solutions
b. taste bitter
c. feel slippery
d. turn red litmus blue
e. neutralize acids

REACTIONS IN SOLUTION - DEMO J1

Purpose:

To classify reactions in solution according to three main reaction types.

Procedure:

Your teacher will demonstrate the following reactions which occur in solution. Identify each reaction according to reaction type, then write a balanced equation for each reaction.

1. zinc + sulfuric acid

   Reaction type: ____________________________

   Balanced equation: ____

2. lead(II) nitrate + potassium iodide

   Reaction type: ____________________________

   Balanced equation: ____

3. calcium + water

   Reaction type: ____________________________

   Balanced equation: ____

4. hydrochloric acid + potassium hydroxide

   Reaction type: ____________________________

   Balanced equation: ____

5. chlorine + sodium bromide

   Reaction type: ____________________________

   Balanced equation: ____

6. barium hydroxide + sulfuric acid

Reaction type: ________________________

Balanced equation: ____

7. bromine + potassium iodide

Reaction type: ________________________

Balanced equation: ____

8. lead nitrate + sodium bromide

Reaction type: ________________________

Balanced equation: ____

9. vinegar + lye

Reaction type: ________________________

Balanced equation: ____

10. chromium(III) nitrate + sodium hydroxide

Reaction type: ________________________

Balanced equation: ____

A. Write balanced equations for each of the following reactions in solution. If the reaction is a precipitation reaction, identify the precipitate with the subscript "s". If the reaction is a neutralization, write "acid", "base", "water" and "salt" under the appropriate chemicals.

1. $HC\ell_{(aq)} + KOH_{(aq)} \longrightarrow$ ______ + ______

2. $BaC\ell_{2(aq)} + Na_2SO_{4(aq)} \longrightarrow$ ______ + ______

3. $NH_4C\ell_{(aq)} + Pb(CH_3COO)_{2(aq)} \longrightarrow$ ______ + ______

4. $LiOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow$ ______ + ______

5. $H_2SO_{3(aq)} + NaOH_{(aq)} \longrightarrow$ ______ + ______

6. $AgNO_{3(aq)} + Li_2CO_{3(aq)} \longrightarrow$ ______ + ______

7. $RbI_{(aq)} + HgNO_{3(aq)} \longrightarrow$ ______ + ______

8. $Ca(OH)_{2(aq)} + HNO_{3(aq)} \longrightarrow$ ______ + ______

9. $Ca(OH)_{2(aq)} + Na_3PO_{4(aq)} \longrightarrow$ ______ + ______

10. $H_3PO_{4(aq)} + Mg(OH)_{2(aq)} \longrightarrow$ ______ + ______

B. In Edmonton, the treatment of raw water taken from the North Saskatchewan River involves chemical processes that include precipitation reactions. For example, the lime-soda ash water softening process involves the removal of bicarbonates and sulfates of calcium and magnesium (which cause hardness in the water) by precipitating these out as calcium carbonate and magnesium hydroxide. Complete the following equations which illustrate the lime-soda ash water softening process.

1. $Ca(HCO_3)_{2(aq)} + Ca(OH)_{2(aq)}$ (slaked lime) $\longrightarrow$ ______ + ______

2. $Mg(HCO_3)_{2(aq)} + Ca(OH)_{2(aq)} \longrightarrow$ ______ + ______

3. $CaSO_{4(aq)} + Na_2CO_{3(aq)}$ (washing soda) $\longrightarrow$ ______ + ______

4. $MgSO_{4(aq)} + Ca(OH)_{2(aq)} \longrightarrow$ ______ + ______

C. Write a balanced equation to show a lab preparation of copper(II) sulfide by a precipitation reaction.

D. Write a balanced equation to show a lab preparation of barium phosphate by a neutralization reaction.

E. Barium sulfate, $BaSO_4$, has low solubility and will precipitate out if solutions containing barium ions and sulfate ions are mixed. Write balanced equations to show the formation of $BaSO_{4(s)}$ by a precipitation reaction and by a neutralization reaction.

REACTIONS IN SOLUTION

NET IONIC EQUATIONS

Writing Ionic Equations

Many chemical reactions can be represented by three different kinds of equations: *nonionic equations*, *total ionic equations* and *net ionic equations*. For reactions in aqueous solution, the most correct are ionic equations since in the ionizing water media, substances that are electrolytes undergo dissociation into ions. The ionic species in aqueous solution subsequently react as ions.

1. Nonionic Equations

In nonionic equations, the elements and compounds are written in their molecular or formula unit forms.
Example:

$$2AgNO_{3(aq)} + BaCl_{2(aq)} \longrightarrow 2AgCl_{(s)} + Ba(NO_3)_{2(aq)}$$

2. Total Ionic Equations

In total ionic equations, elements and compounds are written in the forms in which they are predominately present: electrolytes* as ions; nonelectrolytes, precipitates and gases in their molecular or formula unit forms.
Example:

$$2Ag^{1-}(aq) + 2NO_3^{1-}(aq) + Ba^{2+}(aq) + 2Cl^-(aq) \longrightarrow 2AgCl_{(s)} + Ba^{2+}(aq) + 2NO_3^-(aq)$$

3. Net Ionic Equations

In net ionic equations, only those molecules, formula units or ions that have changed (predominant reacting species) are included in the equation; ions or molecules that do not change (spectator species) are omitted.
Example:

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

*Essentially strong electrolytes are written as ions and weak electrolytes are written in their molecular form. However, such distinction is relevant only when talking about acids and will be made when acids and bases are considered.

REACTIONS IN SOLUTION
NET IONIC EQUATIONS

It should be remembered that in both total and net ionic equations charge as well as atoms have to be balanced. Because ions are electrically charged, chemical equations are often not neutral in charge and end up with a net electrical charge. The net electrical charge of an ionic equation as well as its atoms should be in balance. Therefore, a balanced ionic equation will have the same net electrical charge on both sides of the equation, whether it is zero, positive or negative.

The following is a summary of rules to observe when writing ionic equations:

1. Electrolytes are written in their ion form.
2. Nonelectrolytes are written in their molecular forms.
3. Insoluble substances, precipitates and gases are written in their molecular or formula unit forms.
4. The net ionic equation include only those substances that have undergone a chemical change, i.e.; predominant reacting species.
5. Equations must be balanced, both in atoms and in electrical charge.

The following examples illustrate the writing of ionic equations.

Example I:

$Mg(s) + 2HCl(aq) \longrightarrow MgCl_{2(aq)} + H_{2(g)}$ (nonionic equation)

$Mg(s) + 2H^+(a) + 2Cl^-(aq) \longrightarrow Mg^{2+}(aq) + 2Cl^-(aq) + H_{2(g)}$ (total ionic equation)

$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_{2(g)}$ (net ionic equation)

Example II:

$Na_2SO_4(aq) + Ba(NO_3)_{2(aq)} \longrightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$ (nonionic equ.)

$2Na^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \longrightarrow BaSO_{4(s)} + 2Na^+(aq) + 2NO_3^-(aq)$ (total ionic equation)

$SO_4^{2-}(aq) + Ba^{2+}(aq) \longrightarrow BaSO_{4(s)}$ (net ionic equation)

Example III:

$HNO_{3(aq)} + KOH(aq) \longrightarrow KNO_{3(aq)} + H_2O(\ell)$ (nonionic equation)

$H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \longrightarrow K^+(aq) + NO_3^-(aq) + H_2O$ (total ionic equation)

$H^+(aq) + OH^-(aq) \longrightarrow H_2O$ (net ionic equation)

REACTIONS IN SOLUTION
NET IONIC EQUATIONS

Write the total ionic equations and the net ionic equation for each of the following reactions. All reactions are in water solution.

1. Sodium metal reacts with water.

2. Lead(II) nitrate reacts with hydrgen sulfide gas to yield a precipitate of lead(II) sulfide.

3. Chloric acid is neutralized with potassium hydroxide.

4. Hydorchloric acid is added to a solution of barium hydroxide.

5. Magnesium metal is added to an aqeous solution of hydrogen bromide.

6. Calcium matal reacts with water.

REACTIONS IN SOLUTION
NET IONIC EQUATIONS

7. Hyrochloric acid is added to marbel chips ($CaCO_3$). One of the products, H_2CO_3, immediately decomposes into H_2O and CO_2.

8. Aqueous solutions of sodium sulfate and barium bromide are mixed.

9. An aqueous solution of washing soda, Na_2CO_3, is added to remove Ca^{2+} ions from water that contains dissolved calcium sulfate.

10. A solution of magnesium hydroxide is used to neutralize a dilute solution of nitric acid.

REACTIONS IN SOLUTION - LAB J1

Purpose: To study the effect of concentration on the extent of a chemical reaction.

Materials:

Part I

1- zinc strip
1- 100 ml graduated cylinder
1- 250 ml beaker
1- balance
1- wash bottle
1- bottle containing $CuSO_4$ solution

Part II

1- bottle containing Na_2SO_4 solution
1- bottle containing $BaCl_2$ solution
1- 18x150 mm test tube and solid rubber stopper
1- test tube rack
1- metric ruler
1- 10 ml graduated cylinder

Part III

1- 50 ml graduated cylinder
1- bottle containing 0.100 M HCl solution
1- dropping bottle of phenolphthalein
1- 250 ml erlenmeyer flask
1- bottle of NaOH solution

Day I:

Part I:

1. Determine the initial mass of a clean dry zinc strip.
2. Use a 100 ml graduated cylinder to obtain 100 ml of one of the four solutions of $CuSO_4$ prepared for this lab. Place the 100 ml of $CuSO_4$ solution in a clean 250 ml beaker.
3. Record the number of the $CuSO_4$ solution used in the data table.
4. Place the zinc strip in the $CuSO_4$ solution.
5. Record in the data table observations of any changes.
6. Initial the beaker and set the beaker aside until next day.

Part II:

1. Use a 10 ml graduated cylinder to obtain 10.0 ml of Na_2SO_4 solution and transfer this solution into an 18x150 mm test tube. (Use a metric ruler to confirm the size of the test tube).
2. Clean the 10 ml graduated cylinder and use it to transfer 10.0 ml of one of the $BaCl_2$ solutions into the same test tube.
3. Record the number of the $BaCl_2$ solution used in your data table.
4. Record in the data table observations of any changes.
5. Stopper the test tube then invert several times.
6. Initial the test tube and allow it to stand overnight in the test tube rack.

Part III.

1. Use the clean 100 ml graduated cylinder to obtain 20 ml M HCl solution from a stock bottle. Transfer the solution into a 250 ml erlenmeyer flask.
2. Add about three drops of phenolphthalein indicator to the solution in the erlenmeyer flask.
3. Use the clean 100 ml graduated cylinder to obtain an initial volume of 50 ml of one of the four NaOH solutions prepared for this lab.
4. Record the number of the NaOH solution used in your data table.
5. While swirling the flask to mix the solutions, slowly add the NaOh solution to the acid A LITTLE AT A TIME until a pink colur just persists.
6. Record the final volume of NaOH solution remaining in the graduated cylinder.
7. Record in the data table observations of any changes.
8. Discard all of the solutions into the sink. Clean all of the equipment.

Day 2:

Part I:

1. Record in the data table any visual changes. Note particularly the color of the solution.
2. Use a wash bottle to rinse the zinc strip. Collect the rinse water in the 250 ml beaker.
3. Use a paper towel to wipe off any residue adhering to the zinc strip. Dry the zinc strip.
4. Determine the final mass of the zinc strip.
5. Decant the liquid from the 250 ml beaker into the sink. Discard the solid residue into a solid waste container.
6. Return the zinc strip as directed by the teacher.

Part II:

1. Use a metric ruler to measure to 0.1 cm the height of precipitate in the test tube. Record in the data table the height of precipitate.
2. Discard the contents of the test tube in the sink. Clean all equipment.

REACTIONS IN SOLUTIONS- LAB J1

Data and Questions:

Part I - Student Data:

1. Number of $CuSO_4$ solution used ____________________
2. Number of balance used ____________________
3. Initial mass of zinc strip ____________________
4. Final mass of zinc strip ____________________
5. Mass of zinc reacted ____________________
6. Observed changes ____________________

Group Data:

	Part I: Mass of Zinc Reacted (g)			
	Solution I	Solution 2	Solution 3	Solution 4
Average				

Questions:

1. Considering that the volume of $CuSO_4$ solution was constant, why did the mass of zinc reacted vary?

2. List the solution numbers in order of concentration from least concentrated to most concentrated.

3. Did the initial solutions of $CuSO_4$ differ in appearance?

4. Write a balanced equation for this reaction.

REACTIONS IN SOLUTION - LAB J1

Data and Questions:

Part II - Student Data:

1. Number of $BaCl_2$ solution used ____________
2. Height of precipitate ____________
3. Observed changes

Group Data:

Part II: Height of Precipitate (cm)				
	Solution 1	Solution 2	Solution 3	Solution 4
Average				

Questions:

1. Considering that the volume of $BaCl_2$ solution was left constant, why did the height of the precipitate vary?

2. List the $BaCl_2$ solution numbers in order of concentration from least concentrated to most concentrated.

3. Did the initial solutions of $BaCl_2$ differ in appearance?

4. Write a balanced equation for this reaction.

Part III - Student Data:

1. Number of NaOH solution used ____________
2. Initial volume of NaOH 50.0 ml
3. Final volume of NaOH ____________
4. Volume of NaOH used ____________
5. Observed changes

Group Data:

	Part III: Volume of NaOH (ml)			
	Solution 1	Solution 2	Solution 3	Solution 4
Average				

Questions:

1. Considering the volume of the HCl solution was left constant, why did the volume of the NaOH solution vary?

2. List the NaOH solution numbers in order of concentration from least concentrated to most concentrated.

3. Did the initial solutions of NaOH differ in appearance?

4. Write a balanced equation for this reaction.

REACTIONS IN SOLUTION

A SINGLE REPLACEMENT REACTION IN SOLUTION - DEMO J2

Purpose: To determine the concentration of a solution of copper (II) sulfate.

Procedure:

Day 1:

1. Measure 100.0 ml of the $CuSO_4$ solution in a graduated cylinder and transfer this solution to a clean 250 ml beaker. Rinse the cylinder into the beaker with distilled water from a wash bottle.
2. Place the zinc into the copper (II) sulfate solution so that over half of the strip is immersed.

Day 2:

3. Add about 10 - 15 ml of 6 M HCl to the reaction beaker. After 1 minute remove the zinc strip. Allow the reaction beaker to stand for 5 - 10 minutes. (The reaction of zinc with the acid loosens any copper adhering to the zinc strip. Further reaction consumes any particles of zinc that may be mixed with the copper. Allow it to stand until no more H_2 bubbles form.)
4. Determine and record the mass of a piece of filter paper to 0.01 g.
5. Filter and wash the contents of the reaction beaker.

NOTE these filtration techniques:

a. folding a filter paper
b. "wetting in" a filter paper
c. pouring into a filter funnel
d. maximum liquid level in a filter funnel
e. placement of funnel tip in a waste beaker
f. rinsing of a reaction beaker into a filter funnel

REACTIONS IN SOLUTION

A SINGLE REPLACEMENT REACTION IN SOLUTION - DEMO J2

6. Remove the filter paper with copper carefully, open the paper, and set aside overnight to dry.

Day 3:

7. Determine and record the mass of the dry filter paper with copper to 0.01 g.

Data:

volume of $CuSO_4$ solution	________
mass of filter paper	________
mass of filter paper and copper	________
mass of copper produced	________

Calculations and Questions: (Show All Work)

1. Write the balanced equation for the reaction.

REACTIONS IN SOLUTION

A SINGLE REPLACEMENT REACTION IN SOLUTION - DEMO J2

2. Calculate the number of moles of Cu produced.

3. Calculate the number of moles of $CuSO_4$ present in the initial solution.

4. Calculate the concentration of $CuSO_4$ in the initial solution.

5. Write a dissociation equation for copper(II) sulfate and find the concentration of cupric ions in the initial solution.

6. The mass loss of the zinc strip cannot be used to calculate the initital concentration of $CuSO_4$ solution using the procedure in this experiment. Explain.

7. What is the balanced equation for the reaction used to loosen the copper particles from the zinc strip?

REACTIONS IN SOLUTION

SOLUTION STOICHIOMETRY

Introduction

For most chemical reactions one or more of the reactants is in solution. Often the ease and accuracy of stoichiometric calculations are increased when they are based on the volume of a solution involved. In general, working with solutions of known concentration makes possible much greater accuracy and ease of handling than gravimetric techniques.

For example:

A one milligram (0.00100 g)sample of NaOH(s) is impossible to measure on a centigram balance. However a 10.0 ml. sample of 0.00250 M NaOH solution is easily measured to high accuracy. This volume contains 1.00 mg NaOH.

Thus calculations involving high accuracy or very small quantities are often based on reactions of solutions.

Solution Stoichiometry Calculations

Example #1:

Problem: If 10.0 ml of 0.020 M HCl react exactly with 12.0 ml of $Ba(OH)_2$ solution, find the molarity of the base solution.

Step I: Write a balanced chemical equation for the reaction.

$$2\ HCl + Ba(OH)_2 \longrightarrow BaCl_2 + 2\ HOH$$

Step II: Calculate # of moles of HCl.

$$\#\ \text{moles HCl} = \#\ \ell \times M$$
$$= 0.0100\ \ell \times 0.020\ \frac{\text{mol}}{\ell}$$
$$= 0.00020\ \text{mol}$$

Step III: Use the mole ratio from the balanced equation to calculate # of moles of $Ba(OH)_2$.

$$\#\ \text{moles } Ba(OH)_2 = \#\ \text{mol HCl} \times \frac{\text{coeff } Ba(OH)_2}{\text{coeff HCl}}$$
$$= 0.00020\ \text{mol} \times \frac{1}{2}$$
$$= 0.00010\ \text{mol}$$

Step IV: Calculate the molarity of the $Ba(OH)_2$ solution.

$$M\ Ba(OH)_2 = \frac{\#\ \text{mol}}{\#\ \ell}$$
$$= \frac{0.00010\ \text{mol}}{0.0120\ \ell}$$
$$= 0.0083\ \frac{\text{mol}}{\ell}$$
$$= 0.0083\ M$$

SOLUTION STOICHIOMETRY

Example #2:

Problem: If 200 ml of 0.100 M $AgNO_3$ completely reacts with copper, what mass of silver will be produced?

Step I: Write the balanced chamical equation for the reaction.

$$Cu + 2\ AgNO_3 \longrightarrow 2\ Ag + Cu(NO_3)_2$$

Step II: Calculate # of moles of $AgNO_3$.

$$\#\ \text{moles}\ AgNO_3 = \#\ell \times M$$
$$= 0.200\ \ell \times 0.100\ \frac{mol}{\ell}$$
$$= 0.0200\ mol$$

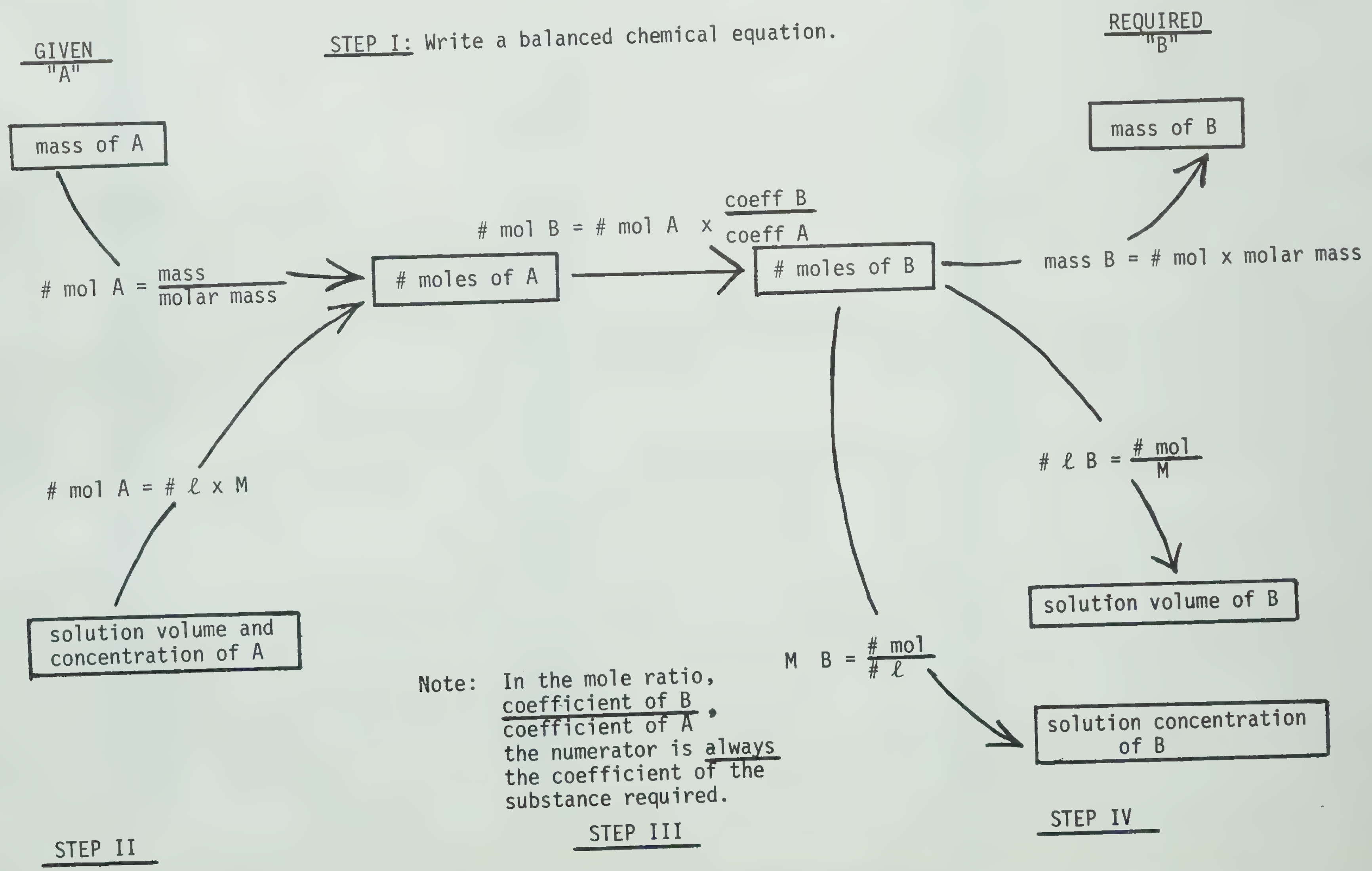
Step III: Use the mole ratio from the balanced equation to calculate the # moles of Ag.

$$\#\ \text{mol Ag} = \#\ \text{mol}\ AgNO_3 \times \frac{\text{coeff Ag}}{\text{coeff}\ AgNO_3}$$
$$= 0.0200\ mol \times \frac{2}{2}$$
$$= 0.0200\ mol$$

Step IV: Calculate the mass of Ag produced.

$$\text{mass of Ag} = \#\ \text{mol} \times \text{molar mass}$$
$$= 0.0200\ mol \times 108\ g/mol$$
$$= 2.16\ g$$

STOICHIOMETRY FLOW CHART



REACTIONS IN SOLUTION

SOLUTION STOICHIOMETRY

Problems

1. If 120 ml of a $Pb(NO_3)_2$ solution is needed to completely precipitate PbI_2 from 80.0 ml of 0.200 M NaI, what is the concentration of the $Pb(NO_3)_2$ solution?

2. What volume of 1.80 M NaOH is needed to neutralize 45.0 ml of 0.800 M H_2SO_4?

3. What volume of 0.500 M sodium chromate is needed to precipitate all the barium ions from 300 ml of a 0.020 M barium chloride solution?

SOLUTION STOICHIOMETRY

4. In order to precipitate all the cobalt (II) ions from 150 ml of 0.500 M cobalt(II) chloride, it was necessary to use 28.4 ml of a sodium hydroxide solution. What was the concentration of the sodium hydroxide solution?

5. A water sample was found to contain a small amount of dissolved ferric nitrate. If 4.8 ml of 0.020 M sodium hydroxide were required to precipitate all the ferric ions from 800 ml of sample, what was the concentration of ferric ions in the sample?

6. A student used 14.40 ml of KOH solution to neutralize 4.80 ml of 0.200 M H_3PO_4. What was the molarity of the base?

REACTIONS IN SOLUTION

GRAVIMETRIC AND SOLUTION STOICHIOMETRY

1. What mass of copper is required to react completely with 250 ml of 0.100 $AgNO_3$ solution?

2. What volume of 2.00 M HCl is needed to neutralize 1.20 g of solid NaOH?

3. A piece of aluminum is placed in a beaker containing 500 ml of H_2SO_4 solution. Using the data table below, calculate the concentration of the H_2SO_4 solution.

initial mass of Al	15.14 g
final mass of Al	9.74 g

REACTIONS IN SOLUTION

GRAVIMETRIC AND SOLUTION STOICHIOMETRY

4. In order to neutralize 3.78 g of solid oxalic acid, $H_2C_2O_4 \cdot 2\ H_2O$, it was necessary to add 125 ml of a lithium hydroxide solution. What was the concentration of the lithium hydroxide solution?

5. What would be the final mass of a 1.20 kg Mg bar after complete reaction with 2.40 ℓ of 6.00 M HCl? (Give your answer in kilograms.)

6. Chlorine gas was bubbled through 120 ml of 0.300 M NaBr until all the bromide ions were replaced. How many moles of chlorine gas reacted?

REACTIONS IN SOLUTION

SOLUTION STOICHIOMETRY - AN OVERVIEW

1. A 100 ml sample of sulfuric acid is completely reacted with zinc. 0.216 moles of hydrogen are produced. Find the molarity of the sulfuric acid.

2. What volume of 3.00 M HNO_3 is needed to neutralize 450 ml of 0.0100 M $Ca(OH)_2$?

3. Chlorine gas is bubbled through 2.50 ℓ of a 0.120 M NaI solution. What is the maximum mass of iodine that could be produced?

REACTIONS IN SOLUTION

SOLUTION STOICHIOMETRY - AN OVERVIEW

4. When a strip of aluminum is placed in 165 ml of 0.240 M nickel(II) nitrate, what would be the concentration of aluminum nitrate in the final solution, assuming all the nickel(II)nitrate reacts and the volume remains constant?

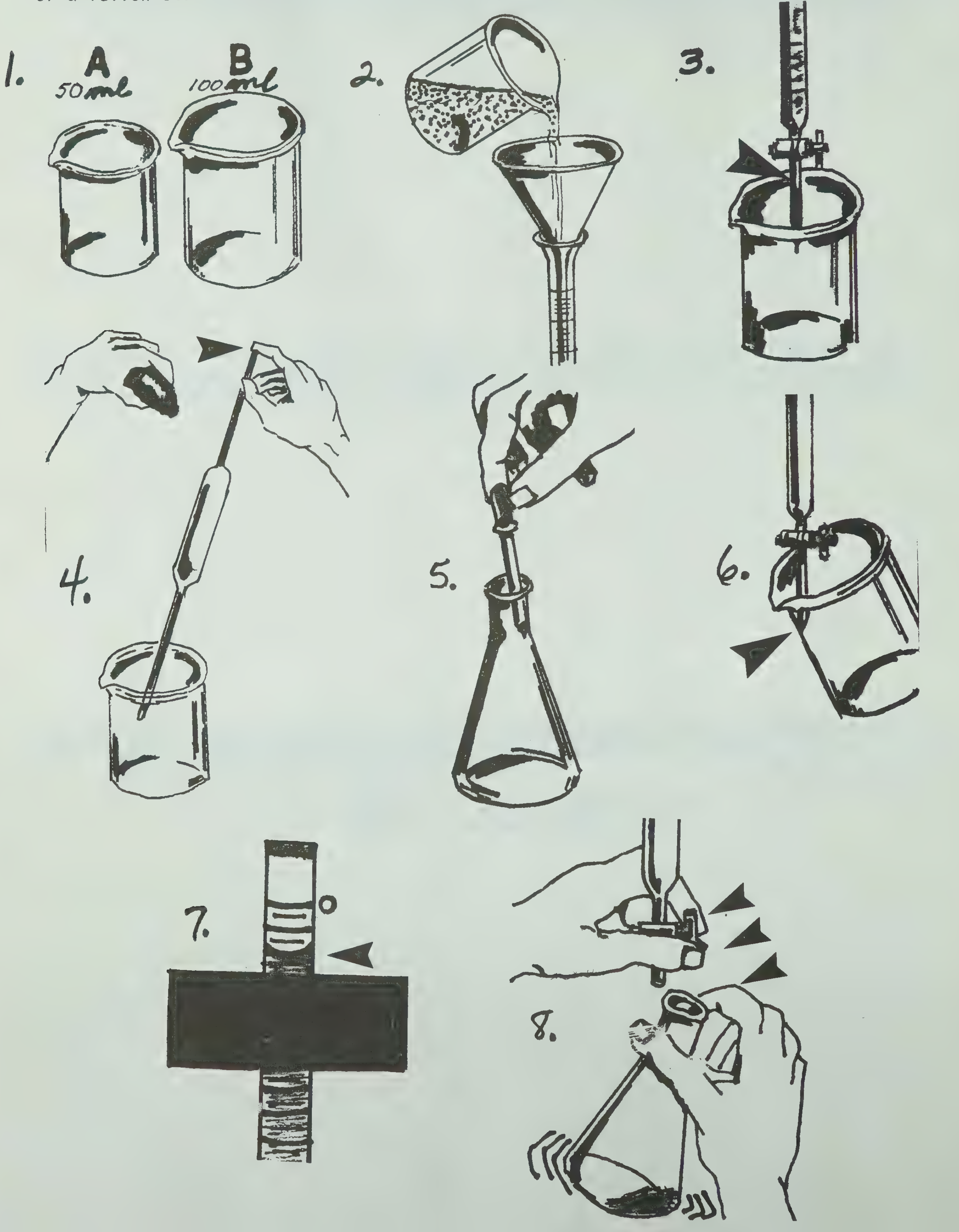
5. What volume of 0.15 M silver nitrate is needed to produce 35.8 g of silver chloride by reacting the silver nitrate with excess sodium chloride solution?

6. The following data were obtained from a titration experiment using 0.250 M KOH and 10.0 ml ml of H_2SO_4 solution of unknown concentration. Determine the concentration of the H_2SO_4 solution.

	base
final burette reading	43.73 ml
initial burette reading	2.48 ml

TITRATION - DEMO J3

The following pictorial presentation of the procedures involved in a titration parallel the steps set out in Titration Lab II - Unit H. Write in notes as the teacher or a fellow student demonstrates.



TITRATION - DEMO J3

Data and Calculation:

Titration of Standardized 0.200 M H_2SO_4 with Aqueous NaOH

Trial Number	Trial #1	Trial #2	Trial #3
Final Buret Reading(ml)			
Initial Buret Reading(ml)			
Volume of NaOH Used (ml)			

1. Calculate the average volume of aqueous NaOH required to neutralize 10.0 ml of 0.200 M $H_2SO_{4(aq)}$. Show your work.

average volume of aqueous NaOH =

2. Use the average volume of aqueous NaOH from above to calculate the concentration of the NaOH solution. Show all steps in the calculation.

concentration of aqueous NaOH =

REACTIONS IN SOLUTION

TITRATION - LAB J2

Purposes:

1. To learn some basic laboratory techniques used in performing titrations.
2. To determine the concentration of a sodium hydroxide solution by titration with a standardized solution of sulfuric acid.
3. To determine the concentration of various acid - base commercial products such as vinegar, soft drinks and household ammonia.

Materials:

- 1 - buret
- 1 - 10 ml pipet
- 1 - pipet bulb
- 1 - 250 ml erlenmeyer flask
- 1 - 400 ml beaker
- 1 - 100 ml beaker
- 1 - 50 ml beaker
- 1 - dropper bottle of phenolphthalein indicator
- 1 - short stem funnel
- 1 - ring stand
- 1 - buret clamp
- 1 - wash bottle with distilled water
- 1 - meniscus finder

Procedure: Refer to the Titration Demonstration notes for a description of proper laboratory techniques.

Day I:

1. a. Wash and dry a 50 ml beaker and label it "A".
   b. Wash and dry a 100 ml beaker and label it "B".
   c. Take the beakers to the stock bottles and obtain about 45 ml of acid, and about 80 ml of the base.
2. a. Place a short stem funnel in the buret and open the buret stopcock.
   b. Rinse the buret with about a 5 ml portion of the base.
   c. Repeat the rinse twice more.
   d. Close the stopcock after the final rinse has drained.
3. a. Fill the buret with the base to a level between the 0 and 5 ml marks.
   b. Open the stopcock fully for an instant to fill the buret tip and to remove any air bubbles from the tip. (Persistent air bubbles may be removed by quickly rotating the stopcock 180^{o}.) (This step is IMPORTANT.)
4. a. Pipet a 10.0 ml sample of distilled water into the 400 ml waste beaker.
   b. Pipet a 10.0 ml sample of acid into the 400 ml waste beaker. (Steps a and b are necessary to wash and rinse the pipet. Exactly 10.0 ml is used in this case for practice only.)
   c. Pipet a 10.0 ml sample of acid into a clean (not necessarily dry) 250 ml erlenmeyer flask.
5. Add about 3 - 4 drops of phenolphthalein indicator to the acid in the erlenmeyer flask. (Note: An indicator is a substance which indicates the end point or completion of a chemical reaction by a change in color.)
6. Remove any drop of base from the buret tip by touching the tip to the side of the 400 ml waste beaker.
7. Record the initial buret reading. (For greater accuracy use a meniscus finder.)
8. a. Place the erlenmeyer flask on a white sheet of paper under the buret and lower the buret to a point just above the flask.
   b. Following proper titration techniques, start adding the base to the erlenmeyer flask.

TITRATION - LAB J2

c. As the pink color begins to linger upon swirling the erlenmeyer flask, control the stopcock carefully to add the base more slowly.
d. At the endpoint remove the final drop from the buret tip.
e. Use a wash bottle to rinse down the sides of the flask with distilled water.
f. Record the final buret reading.
g. Discard the contents of the erlenmeyer flask and rinse at least twice with distilled water.

9. Do the titration (step 4c to 8g) two more times.

10. Clean all the glassware. Rinse the buret with distilled water and leave the buret INVERTED in its clamp with the stopcock OPEN.

Data and Claculations:

Day 1: Titration of Standardized 0.0300 H_2SO_4 with Aqueous NaOH

Trial Number	Trial #1	Trial #2	Trial #3
Final Buret Reading(ml)			
Initial Buret Reading(ml)			
Volume of Base Used(ml)			

1. Calculate the average volume of base required to neutralize 10.0 ml of 0.0300 M H_2SO_4.

2. Use the average volume of aqueous NaOH from above to calculate the concentration of the base solution. Show all steps in the calculations.

concentration of aqueous NaOH =

TITRATION - LAB J2

3. Why must the 50 ml and the 100 ml beaker be initially dry while the erlenmeyer flask can be initially wet?

4. Teacher Comments on Laboratory Technique and Results

Procedure:

Day 2:

1. Follow the procedure outline for Day 1. Perform only a SINGLE titration for each of the following commercial products. Take only enough reagent from the stock bottles to complete ONE titration.

   a. VINEGAR titration:

   Pipet 10.0 ml of one of the available brands of vinegar into an erlenmeyer flask. Rinse and fill the buret with standardized NaOH solution from Day 1, and titrate. Record data in table provided.

   b. POP titration:

   Pipet 10.0 ml of one of the available brands of soda pop in an erlenmeyer flask. Titrate with the standardized NaOH solution. Record data.

   c. HOUSEHOLD AMMONIA titration:

   Pipet 10.0 ml of standardized H_2SO_4 solution (from Day 1) into an erlenmeyer flask. Rinse the buret several times and fill with one of the available brands of household ammonia. Titrate. Record data.

2. Record all class data from the master data table provided by your teacher (or as directed).

Data and Calculations:

Day 2:

Single Group Data

Titration	Vinegar with NaOH	Pop with NaOH	H_2SO_4 with Ammonia
Commercial Brand Name			
Final Buret Reading(ml)			
Initial Buret Reading(ml)			
Volume of Base Used(ml)			

REACTIONS IN SOLUTION

TITRATION - LAB J2

Class Data

Titration	Vinegar with NaOH			Pop with NaOH			H_2SO_4 with Ammonia		
Brand Name									
Volume of Base Used									
Average Volume of Base Used									
Concentration of Commercial Product									
*Corrected Concentration									

* The vinegar has been diluted 10 fold. (Multiply your calculated concentration by 10.) The household ammonia has been diluted 20 fold. (Multiply your calculated concentration by 20.)

Question:

Write a balanced equation for each of the three neutralizations. Assume that vinegar is aqueous CH_3COOH, pop is aqueous H_2CO_3 and ammonia is aqueous NH_4OH.

REACTIONS IN SOLUTION

TITRATION PROBLEMS

Show All Work

1. For a titration 10.0 ml of 0.150 M phosphoric acid (H_3PO_4) was pipetted into a 250 ml erlenmeyer flask. Titrating to an endpoint with potassium hydroxide provided the following data:

final buret reading ...... 34.5 ml
initial buret reading .... 2.3 ml

What was the concentration of the base solution?

2. A titration between standardized 0.400 M NaOH and an aqueous solution of H_2CO_3 provided the following data.

	Base	Acid
final buret reading	14.7 ml	23.4 ml
initial buret reading	2.6 ml	0.8 ml

What was the molarity of the acid solution?

3. A titration to determine the amount of iron in iron ore involves the following chemical reaction.

$$10\ FeSO_4 + 2\ KMnO_4 + 8\ H_2SO_4 \longrightarrow 5\ Fe_2(SO_4)_3 + K_2SO_4 + 2\ MnSO_4 + 8\ H_2O$$

If 10.0 ml of aqueous $FeSO_4$ was titrated with standardized 0.100 M $KMnO_4$ to obtain the following data, what was the concentration of the $FeSO_4$?

final buret reading ...... 42.6 ml
initial buret reading .... 7.3 ml

REACTIONS IN SOLUTION
THE DETERMINATION OF THE UNKNOWN CONCENTRATION
OF A SOLUTION BY GRAVIMETRIC ANALYSIS - LAB J3

Purposes:

1. To determine the concentration of a solution of $Pb(NO_3)_2$.
2. To develop technique in the process of filtration.

Materials:

1 - 10 ml pipet
1 - 100 ml graduated cylinder
1 - centigram balance
1 - filter paper
1 - 50 ml beaker
2 - 250 ml beakers
1 - 400 ml beaker
1 - stirring rod
1 - filter funnel
1 - filter holder
1 - ring stand
1 - wash bottle of distilled water
1 - large watch glass
stock solution of 0.125 M KIO_3
stock solution of aqueous $Pb(NO_3)_2$

Procedure:

(All masses should be obtained to 0.01 g.)

Day 1:

1. Obtain about 25 ml of $Pb(NO_3)_2$ solution in a 50 ml beaker.
2. Pipet 10.0 ml of the lead (II) nitrate solution of unknown concentration into a clean 250 ml beaker.
3. Obtain about 125 ml of 0.125 M KIO_3 in a 250 ml beaker.
4. Using a graduated cylinder, transfer 100 ml of 0.125 M KIO_3 solution to the beaker containing the $Pb(NO_3)_2$ solution.
5. Record the number of the centigram balance used.
6. Determine and record the mass of a piece of filter paper.
7. Set up a filtration apparatus.
8. Filter and wash the precipitate, following the procedure demonstrated previously.
9. When all the liquid has drained, carefully remove the filter paper. (Remember how easily wet paper tears.) Unfold the filter paper so it can dry more easily and place it onto a labelled watchglass to dry overnight.
10. Clean all glassware and the working area.

Day 2:

11. Determine and record the mass of the filter paper plus lead (II) iodate.
12. Clean your working area well. (Discard the precipitate and filter paper into a trash can.)

REACTIONS IN SOLUTION

THE DETERMINATION OF THE UNKNOWN CONCENTRATION OF A SOLUTION BY GRAVIMETRIC ANALYSIS - LAB J3

Data:

Day 1:

balance number ____________

mass of filter paper = ____________

Description of precipitate:

Day 2:

mass of filter paper + $Pb(IO_3)_2$ = ____________

mass of $Pb(IO_3)_2$ = ____________

Calculations and Questions:
(Show All Work)

1. Write the balanced equation for the reaction.

2. Calculate the number of moles of lead (II) iodate precipitated.

3. Calculate the number of moles of lead (II) nitrate in the initial solution sample.

REACTIONS IN SOLUTION
THE DETERMINATION OF THE UNKNOWN CONCENTRATION OF A SOLUTION BY GRAVIMETRIC ANALYSIS - LAB J3

4. Calculate the concentration of $Pb(NO_3)_2$ in the initial solution.

5. (Optional)

   a. Based on the number of moles of $Pb(IO_3)_2$ precipitated (see question 2), calculate the number of moles of KIO_3 that reacted.

   b. Calculate the number of moles of KIO_3 in the initial 100 ml of 0.125 M solution.

   c. Calculate the number of moles of KIO_3 in excess.

6. (Optional) Which reactant limits the amount of precipitate formed? (Which reactant is the "limiting factor" in this reaction?)

7. (Optional) Can the mass of precipitate be predicted from the number of moles of KIO_3 present in the initial 100 ml of 0.125 M KIO_3? Explain your answer.

8. (Optional) What ions are present in the filtrate?

[illegible] BY GRAVIMETRIC ANALYSIS LAB 13

[illegible]culate the concentration of $Pb(NO_3)_2$ in the initial solution.

[illegible] number of moles of KIO_3 [illegible]

DATE DUE SLIP

JUL 27 RETURN

[redacted]

RETURNED APR 27 '78

OCT 6 RETURN

DUE Educ APR 30 '81

JAN 21 RETURN

EDUC MAR 26 '83

RETURN MAR 14 '83

DUE EDUC JUN 27 '83

JUN 27 RETURN

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